CHM129

Periodic Properties

1) Describe the trends for atomic radii as we go down/across the periodic table. Explain why they follow this order. Which of the series below is arranged in the appropriate order:

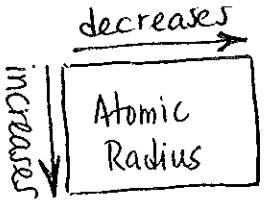
C < N < O < F

F < S < Al < Rb ← correct

Sr < Ca < Mg < Be

Li < Na < K < Ca ← incorrect

incorrect (opposite order)

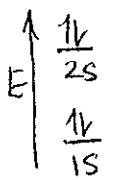


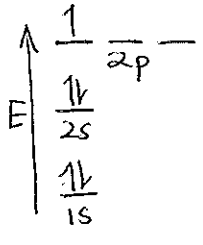
Period (Zeff dominates): e^- are added to the same outer level so shielding by core e^- does not change (same # core e^-). Zeff increases & outer e^- are pulled closer. Atomic size decreases.

Group (n dominates): as we move down, one more shell of inner e^- is added which shield outer e^- effectively. Atomic size increases.

2) Describe the trends for first ionization energies as we go down/across the periodic table. Explain why they follow this order. If you look closely at the numbers on the screen, you will see that the first ionization energy for Be is higher than that for B. The same is true for N and O. Explain this.

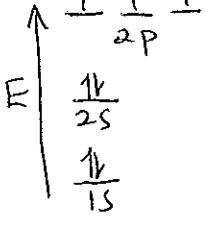
Be > B

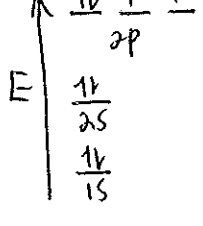




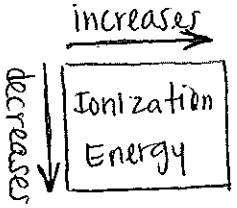
The 3rd valence e^- for B must occupy the 2p subshell, which is empty for Be (higher in energy).

N > O





N has a half-filled shell and O has one more e^- than N (two e^- are paired up). That increases electron-electron repulsions so ionization energy decreases



3) Why alkali metals lose one electron to form +1 cations? On Monday, we observed that K reacts more violently with water that Na. Why (what property is most responsible for this)?

Alkali metals tend to lose e^- so they can attain a stable electron configuration (like a noble gas).

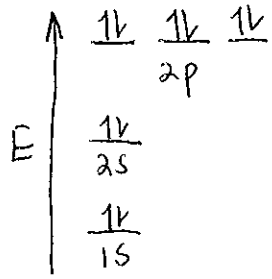
Na: $1s^2 2s^2 2p^6 3s^1$ $\quad Na^+$: $1s^2 2s^2 2p^6$ = [Ne]

[Ne]$3s^1$ ↑ e^- lost

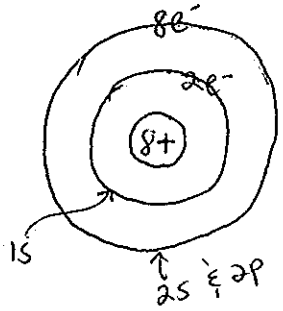
Alkali metal have the lowest ionization energies of any period. This reflects the relative ease with which its outer s electron can be removed. K has a lower ionization energy than Na so it'll lose its outer electron more readily than Na.

Alkali Metals ⇒ very low ionization energies & small negative electron affinity ⇒ form +1 cations.

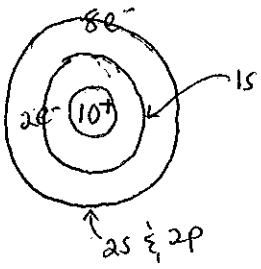
4) The oxide ion, O^{2-}, is isoelectronic (i.e. has exactly the same number and configuration of electrons) with Ne, and yet O^{2-} is bigger than Ne. Explain why.



O^{2-} and Ne



O^{2-}



Ne

$Z_{eff} = Z - S$

$Z_{eff} = 8 - 2 = \underline{6^+}$

$Z_{eff} = 10 - 2 = \underline{8^+}$

O^{2-} and Ne have the same number of electrons but Ne has a larger effective nuclear charge (more protons & same # of core e^-) so it holds the electrons more tightly.

5) Which pair of elements is correctly arranged in order of increasing electron affinity (in terms of absolute numbers)?

(As<Br)

Cl<S Si<P O<B

incorrect